QD
41
A3312
1976
UNIT
I
C. 3

CURR

Ex libris
universitatis
albertaensis

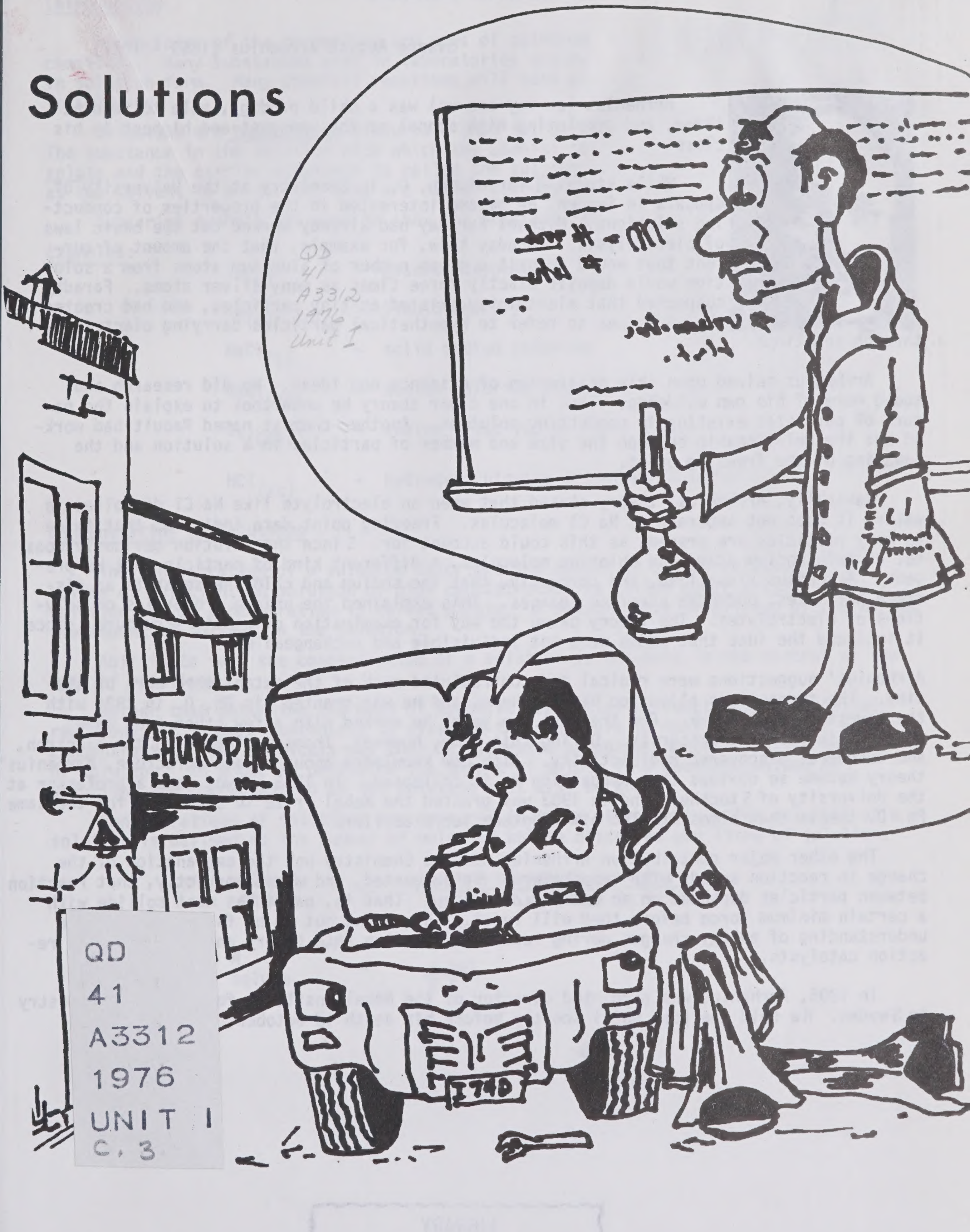
QUAECUMQUE VERA

Digitized by the Internet Archive
in 2021 with funding from
University of Alberta Libraries

https://archive.org/details/alchem20unitijenk

UNIT I

Solutions

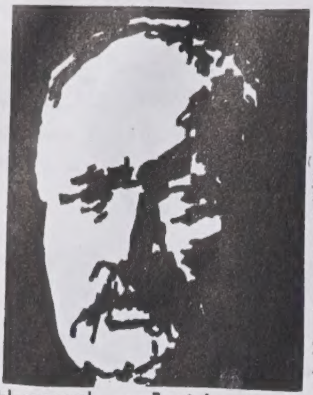

m = # mol
# volume in lit.
CHUKSPIKE
I74D

QD
41
A3312
1976
Unit I

QD
41
A3312
1976
UNIT I
C. 3

SOLUTION CHEMISTRY
ANTE AUGUST ARRHENIUS

Svante August Arrhenius (1859 - 1927)



Arrhenius (ar-ray-nee-us) was a child prodigy, able to read at three, and completing high school as the youngest and highest in his class

While studying for his Ph. D. in Chemistry at the University of Uppsala in Sweden, he became interested in the properties of conducting solutions. Michael Faraday had already worked out the basic laws of electrolysis. Faraday knew, for example, that the amount of current that would deposit a given number of aluminum atoms from a solution would deposit exactly three times as many silver atoms. Faraday suspected that electricity existed as tiny particles, and had created the word *ions* to refer to hypothetical particles carrying electricity through solutions.

Arrhenius seized upon this scattering of evidence and ideas. He did research and added more of his own evidence. Then in one clear theory he undertook to explain the nature of particles existing in conducting solution. Another chemist named Raoult had worked out the relationship between the size and number of particles in a solution and the lowering of the freezing point.

Basically, Arrhenius' theory stated that when an electrolyte like Na Cl dissolves in water, it does not separate as Na Cl molecules. Freezing point data indicated that twice as many particles are present as this could account for. Since the solution obviously does not contain sodium atoms and chlorine molecules, a different kind of particle must be present. Arrhenius visualized, and correctly, that the sodium and chlorine separate as discrete particles, *but with electric charges*. This explained the unique properties of solutions of electrolytes. The theory paved the way for examination of atomic structure, since it included the idea that atoms were not indivisible and unchangeable.

Arrhenius' suggestions were radical and contradicted much of the established work of his time. The suggestions alienated his teachers, and he was granted his Ph. D. in 1884 with the lowest possible mark. For the next ten years he worked with a few other "new wave" chemists in relative obscurity. During this time, however, Thomson discovered the electron, and Becquerel discovered radioactivity. With new knowledge about atomic structure, Arrhenius' theory became so obvious that opposition to it collapsed. In 1895 he was made a professor at the University of Stockholm and in 1903 was granted the Nobel Prize in chemistry for the same Ph. D. thesis that almost failed him nineteen years earlier.

The other major contribution Arrhenius made to Chemistry was the explanation of the change in reaction speeds with temperature. He suggested, and again correctly, that reaction between particles depended on an *activation energy*. That is, particles must collide with a certain minimum force before they will react. This turns out to be fundamental to the understanding of evergy changes during reaction, reaction rate theory and the action of reaction catalysts.

In 1905, Arrhenius was appointed director of the Nobel Institute for Physical Chemistry in Sweden. He held the post until shortly before his death in October, 1927.

LIBRARY
UNIVERSITY OF ALBERTA

SOLUTIONS

Introduction

A knowledge of the properties and uses of solutions is important in the study of chemistry. Many substances used in laboratories are much easier to store and to use in solution form. Many chemical reactions will take place only when the reactants are in solution.

Solutions are homogeneous mixtures - that is, solutions are uniform throughout. The substance in the solution with which the chemist is most concerned is called the solute and the carrier substance is called the solvent. Most common solutions use water as the solvent in which case they are called aqueous solutions.

Subscript symbols are used to show the physical state of a substance.

Examples:

$O_{2(g)}$ - gaseous oxygen

$Hg_{(\ell)}$ - liquid mercury

$NaCl_{(s)}$ - solid sodium chloride

$NaCl_{(aq)}$ - sodium chloride dissolved in water

$I_{2(CCl_4)}$ - iodine dissolved in carbon tetrachloride

$HCl_{(aq)}$ - hydrogen chloride dissolved in water

Concentration of Solutions

The concentration of a solution describes the amount of solute relative to the volume of solution. A solution of high concentration is called concentrated, while a solution of low concentration is called dilute. A solution may be diluted by increasing the amount of solvent in the solution.

Ability to vary the concentration of a solution is valuable in the control of the rate of reactions.

Rather than determine the mass of specific quantities of reactants for chemical reactions, it is more convenient to dissolve the chemicals to make solutions of known concentration. The chemicals may then be dispensed by measuring out specific volumes of solution. In order to determine the amount of solute in a measured volume of solution, the concentration of the solution must also be known.

Concentrations of solutions are measured in terms of molarity. The molarity of a solution is defined as the number of moles of solute dissolved per litre of solution.

$$\text{molarity}_{\text{solute}} = \frac{\text{amount of solute in moles}}{\text{volume of solution in litres}}$$

or

$$M_{\text{solute}} = \frac{\#\ \text{mol}}{\#\ \ell}$$

SOLUTIONS
CONCENTRATION OF SOLUTIONS

The units of molarity are "moles per litre" which is symbolized as "mol/ℓ" or "M" and read as 'molar'. A 0.35M solution would contain 0.35 moles of solute in each litre of solution. Note that concentration is independent of the actual volume of solution.

Solution Problems

1. Calculating Concentration from Moles and Volume

Example 1: If 0.900 mol of NaCl is dissolved to give 500 ml of solution, what is its concentration?

$$\underset{NaCl}{M} = \frac{\#\ mol}{\#\ \ell} = \frac{0.900\ mol}{0.500\ \ell} = 1.80\ M$$

2. Calculating Concentration from Mass and Volume

Example 2: If 17.0 g of $NaNO_3$ is dissolved to give 400 ml of solution, what is its concentration?

$$\#\ moles = \frac{mass}{molar\ mass} = \frac{17.0g}{85.0\ g/mol} = 0.200\ mol$$

$$\underset{NaNO_3}{M} = \frac{\#\ mol}{\#\ \ell} = \frac{0.200\ mol}{0.400\ \ell} = 0.500\ M$$

3. Calculating Moles from Volume and Concentration

Example 3: How many moles of $AlCl_3$ are in 600 ml of a 0.0250M solution?

$$\underset{AlCl_3}{\#\ mol} = M \times \#\ \ell = 0.0250\ mol/\ell \times 0.600\ \ell = 0.0150\ mol$$

4. Calculating Mass from Volume and Concentration

Example 4: What mass of $Mg(CH_3COO)_2$ is necessary to make 400 ml of a 0.0500 M solution?

$$\underset{Mg(CH_3COO)_2}{\#\ mol} = M \times \#\ \ell = 0.0500\ mol/\ell \times 0.400\ \ell = 0.0200\ mol$$

$$\underset{Mg(CH_3COO)_2}{mass} = \#\ moles \times molar\ mass = 0.0200\ mol \times 142\ g/mol = 2.84\ g$$

5. Calculating Volume from Moles and Concentration

Example 5: What volume of 0.750 M $Na_2Cr_2O_7$ contains 1.50 mol of the solute?

$$\underset{Na_2Cr_2O_7}{\#\ \ell} = \frac{\#\ mol}{M} = \frac{1.50\ mol}{0.750\ mol/\ell} = 2.00\ \ell$$

6. Calculating Volume from Mass and Concentration

Example 6: What volume of 0.0200 M $KMnO_4$ contains 0.790 g of solute?

$$\underset{KMnO_4}{\#\ moles} = \frac{mass}{molar\ mass} = \frac{0.790\ g}{158\ g/mol} = 0.00500\ mol$$

$$\underset{KMnO_4}{\#\ \ell} = \frac{\#\ mol}{M} = \frac{0.00500\ mol}{0.0200\ mol/\ell} = 0.250\ \ell \text{ or } 250\ ml$$

Preparation of Solutions

The correct procedure for the preparation of solutions involves the following steps:

a. Calculate the mass of solute required.
b. Dissolve the solute in less than the final volume of solvent.
c. Bring the solution up to final volume.

SOLUTIONS
PREPARATION OF A SOLUTION - DEMO I1

Purpose: To prepare 100.0 ml of a 0.03500M solution of $K_2Cr_2O_7$.

Procedure: 1. a. Calculate the number of moles of $K_2Cr_2O_7$ required:

b. Calculate the mass of $K_2Cr_2O_7$ required.

2. Measure out the required mass of $K_2Cr_2O_7$ and dissolve it in less than 100 ml of H_2O.

a. Why is the mass of the beaker determined?

b. Why is the required mass of $K_2Cr_2O_7$ added to the mass of the beaker on the balance riders (or dial)?

c. How can you tell if too much $K_2Cr_2O_7$ is added to the beaker?

d. Why is distilled water used?

e. Why is the stirring rod rinsed off into the beaker?

SOLUTIONS
PREPARATION OF A SOLUTION - DEMO I1

f. Why is less than 100 ml of water used initially?

3. Bring the volume up to 100.0 ml in a graduated cylinder or volumetric flask.

   a. Why was it unimportant to know the volume of water the $K_2Cr_2O_7$ was dissolved in initially?

   b. Use a diagram to show where the meniscus should be relative to the 100.0 ml mark on the graduated cylinder.

   c. Why is the solution stoppered and inverted several times?

   d. Why should solutions be stored in sealed containers?

SOLUTIONS

SOLUTION PROBLEMS

CONCENTRATION FROM MOLES AND VOLUME

1. What is the concentration of a solution in which 5.00 mol of KCl are dissolved in water to make 2.00 ℓ of solution?

2. What is the molarity of a solution in which 0.240 mol of washing soda, $Na_2CO_3 \cdot 10H_2O$, is dissolved in water to make 480 ml of solution?

3. When 0.300 mol of zinc sulfate is dissolved in water to give 250 ml of solution, what is the concentration of this solution?

4. If 840 ml of a calcium chloride solution contains 0.210 mol of calcium chloride, what is the molarity of this solution?

5. What is the concentration of a solution made by dissolving 0.320 moles of Chile saltpetre (sodium nitrate) to form 1.60 litres of solution?

SOLUTIONS
SOLUTION PROBLEMS

CONCENTRATION FROM MASS AND VOLUME

1. If 3.42g of solid $Ba(OH)_2$ is dissolved in water to make 500 ml of solution, what is the molarity of this solution?

2. When 15.0g of glacial acetic acid is dissolved in water to make 50.0 ml of solution, what is the concentration of the solution?

3. What is the molarity of 500 ml of a solution that contains 12.7g of swimming pool chlorinater $Ca(OCl)_2$?

4. What is the concentration of a solution of baking soda (sodium bicarbonate) which contains 69.2g of baking soda dissolved in water to make 400 ml of solution?

SOLUTIONS

SOLUTION PROBLEMS

5. Various treatment methods are applied to the "raw water" taken from the North Saskatchewan River in Edmonton before general distribution to consumers. Some of the treatment methods involve the addition of various chemicals to the "raw water". The exercise following involves the calculation of the molar concentration of the various chemicals at the time these are added. Show method of calculation.

Chemical Formula	Grams/Litre	Molar Concentration
1. CaO	0.100	
2. Na_2CO_3	0.00667	
3. $A\ell_2(SO_4)_3$	0.00750	
4. $(NH_4)_2SO_4$	0.00300	
5. $C\ell_2$	0.000500	
6. $NaOC\ell$	0.000500	
7. H_2SiF_6	0.00080	

SOLUTIONS
SOLUTION PROBLEMS

MOLES FROM VOLUME AND CONCENTRATION

1. How many moles of solid sodium sulfide are needed to make 3.00ℓ of 0.200 M solution?

2. How many moles of solute are required to prepare 50.0 ml of 2.00 M solution?

3. How many moles of NaOH are contained in 0.550 ml of a 2.50 M NaOH solution?

4. Find the number of moles of aluminum sulfate in 400 ml of 1.20 M $Al_2(SO_4)_3$.

5. How many moles ofcalcium chloride are needed to prepare 50.0 ml of a 0.100 M solution?

SOLUTIONS

SOLUTION PROBLEMS

MASS FROM VOLUME AND CONCENTRATION

1. What mass of sodium chloride is needed to prepare 5.00 ℓ of a 0.100 M solution?

2. How many grams of slaked lime, $Ca(OH)_2$, are contained in 250 ml of a 0.0500 M solution of slaked lime?

3. What mass of solute is required to prepare 750 ml of 2.00 M KOH?

4. Describe how you would prepare 100 ml of a 0.120 M KBr aqueous solution. Show your calculations.

SOLUTIONS
SOLUTION PROBLEMS

VOLUME FROM MOLES AND CONCENTRATION

1. What volume of 3.50 M KNO_3 solution would contain 0.350 mol of solute?

2. What volume of 2.50 M solution would contain 0.140 mol of calcium bromide?

3. What volume of 0.320 M ammonium sulfate solution would contain 0.240 mol of ammonium sulfate?

4. What volume of 0.100 M sodium sulfate solution would contain 3.60 mol of sodium sulfate?

5. Describe how you would prepare a 2.00 molar solution of magnesium sulfate using 3.00 moles of solid magnesium sulfate.

SOLUTIONS
SOLUTION PROBLEMS

VOLUME FROM MASS AND CONCENTRATION

1. How many litres of 0.200 M solution would contain 64.0 g of copper (II) sulfate?

2. What volume of a 0.400 M solution may be prepared from 83.0g of solid potassium iodide?

3. What volume of 0.200 M solution contains 26.0g of ferric sulfate?

4. Describe how you would prepare a 0.250 M solution of KCl using 37.3 g of solid potassium chloride. Show your calculations.

SOLUTION PROBLEMS - AN OVERVIEW

1. What is the concentration of a solution which has 0.240 mol of sodium phosphate dissolved to make 300 ml of solution?

2. Find the number of moles of ammonium sulfite in 80 ml of a 0.14 M solution.

3. What volume of 2.5 M solution would contain 2.0 mol of solute?

4. What is the concentration of a solution in which 43.5g of potassium sulfate is dissolved to yield 500 ml of solution?

SOLUTIONS

SOLUTION PROBLEMS - AN OVERVIEW

5. What volume of a 0.10 M solution may be prepared from 55.5g of calcium chloride?

6. What mass of barium hydroxide is required to prepare 400 ml of a 0.200 M solution? Describe how to prepare the solution.

7. Describe in detail how you would prepare 900 ml of a 0.450 M solution of ammonium dichromate.

SOLUTIONS

DILUTION

Most solutions in labs are stored in a concentrated form. The solutions are then diluted to yield solutions of known concentration as required.

When a solution is diluted, only the amount of solvent is increased. Therefore, the number of moles of solute in the initial concentrated solution is equal to the number of moles of solute in the final diluted solution.

$$\# \text{ moles}_i = \# \text{ moles}_f$$

i = initial
f = final

Since # mol = M x #ℓ, then:

$$M_i \times \#\ell_i = M_f \times \#\ell_f$$

This equation may be used to solve the three types of dilution problems.

1. Calculating Final Concentration

Example 1: What is the concentration of a NaF solution made by diluting 100 ml of a 0.500 M solution to 4.00ℓ?

$$M_i \times \#\ell_i = M_f \times \#\ell_f$$

$$0.500 \text{ mol}/\ell \times 0.100\ell = M_f \times 4.00\ell$$

$$M_f = \frac{0.500 \text{ mol}/\ell \times 0.100\ \ell}{4.00\ell} = 0.0125 \text{ M}$$

2. Calculating Final Volume

Example 2: To what final volume must 10.0 ml of 6.00 M H_2SO_4 be diluted to give a 0.0200M solution?

$$M_i \times \#\ell_i = M_f \times \#\ell_f$$

$$6.00 \text{ mol}/\ell \times 0.0100\ell = 0.0200 \text{ mol}/\ell \times \#\ell_f$$

$$\#\ell_f = \frac{6.00 \text{ mol}/\ell \times 0.0100\ell}{0.0200 \text{ mol}/\ell} = 3.00\ell$$

3. Calculating Initial Volume

Example 3: What volume of 0.500 M $K_2Cr_2O_7$ is necessary to yield 600 ml of a 0.400 M solution?

$$M_i \times \#\ell_i = M_f \times \#\ell_f$$

$$0.500 \text{ mol}/\ell \times \#\ell_i = 0.400 \text{ mol}/\ell \times 0.600\ \ell$$

$$\#\ell_i = \frac{0.400 \text{ mol}/\ell \times 0.600\ \ell}{0.500 \text{ mol}/\ell} = 0.480\ \ell \text{ or } 480 \text{ ml}$$

4. Calculating Concentration When Two Solutions of Different Concentration Are Mixed

Example 4: What is the final concentration when 200 ml of 0.500 M H_2SO_4 and 300 ml of 6.00 M H_2SO_4 are mixed?

$$\#\text{mol}_i = \#\text{mol}_f$$

$$0.500 \frac{\text{mol}}{\ell} \times 0.200\ell + 6.00 \frac{\text{mol}}{\ell} \times 0.300\ell = M_f \times 0.500\ \ell$$

$$M_f = \frac{1.90 \text{ mol}}{0.500\ell} = \frac{3.80 \text{ mol}}{\ell} \text{ or } 3.80 \text{ M}$$

SOLUTIONS
DILUTION PROBLEMS

1. To what volume must 300 ml of 4.00 M C_2H_5OH be diluted to make a 3.00 M solution?

2. What would be the final concentration if 600 ml of 6.00 M HCl were diluted to a volume of 800 ml?

3. What volume of 2.50 M $BaCl_2$ solution is necessary to make 8.00 ℓ of a 0.500 M solution of $BaCl_2$?

4. If 70.0 ml of 0.800 M $ZnBr_2$ is used to make a 0.300 M solution, what is the final volume?

5. What volume of 2.00 M $Cu(NO_3)_2$ is necessary to make 50.0 ml of a 1.70 M solution of $Cu(NO_3)_2$?

SOLUTIONS
DILUTION PROBLEMS

6. If water were added to 500 ml of 3.00 M $FeSO_4$ to give a final volume of 580 ml, what would be the final concentration?

7. If 80.0ℓ of 3.00 M NaOH were released from an industrial plant into a holding tank, what final volume of solution is necessary to yield a final concentration of 0.00100 M?

8. If 60.0ℓ of a 2.50 M toxic substance were poured into a pond to give a final volume of 5.00×10^3 ℓ, what would be the final concentration?

9. To what final volume must the 60.0ℓ of 2.50 M solution in Problem 8 be diluted to make the final concentration 1.00×10^{-6} M (a fairly safe concentration for most toxic substances)?

10. To what volume must 600 ml of a 2.00 M solution of $Ca(BrO_3)_2$ be evaporated to make the solution 4.50 M?

SOLUTIONS

DILUTION PROBLEMS

11. Commercial stock sulfuric acid is 18.0 M. If 600 ml of 3.00 M H_2SO_4 is required, how many ml of stock sulfuric acid and how many ml of water must be mixed.

12. Calculate the final concentration of a solution formed from mixing 1.00 litre of 0.100 M NaOH solution with 1.00 litre of 0.020 M NaOH solution. (NaOH, also known as caustic soda or lye, has widespread use as a cleaner and disinfectant. It should be remembered that NaOH and its solution is very caustic.)

13. What would be the final concentration of $NaHSO_4$ in a solution formed by mixing 400 ml of 0.0500 M $NaHSO_4$ and 600 ml of 0.0200 M $NaHSO_4$? (Soduim bisulfate, $NaHSO_4$, the chief ingredient of many toilet bowl cleaners, depends for its cleaning action on the formation of sulfuric acid on contact with water.)

14. Given a container of 0.750 M Na_3PO_4 and a container of 3.00 M Na_3PO_4. Using only these two solutions (no additional water) how much of each solution must be used to make 120 ml of 1.50 M Na_3PO_4.
(${Na_3PO_4}$, sodium phosphate, is versatile in its usefulness as a scouring powder and as a paint remover.)

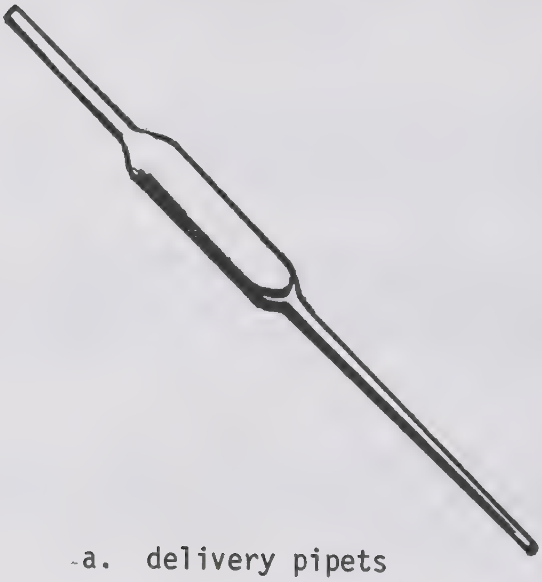
SOLUTIONS

PIPETTING TECHNIQUE - DEMO I2

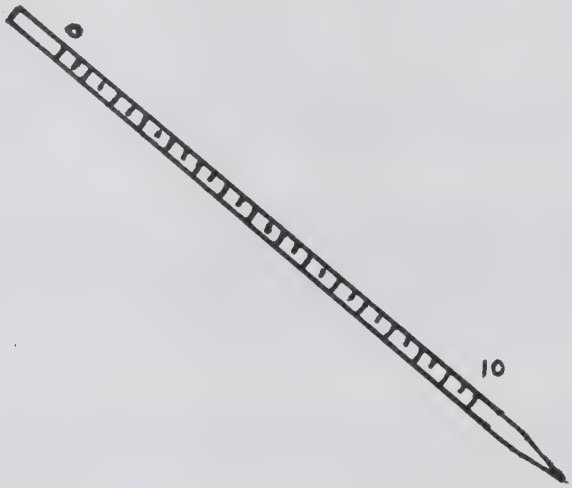
Pipetting is a common technique for measuring out a particular volume of a liquid. A pipet measures a small volume (usually 25 ml or less) to high accuracy. Two basic types of pipets are:

a. delivery pipets, which measure a specific volume only, and

b. graduated pipets, which have a scale and measure incremental volumes.



a. delivery pipets

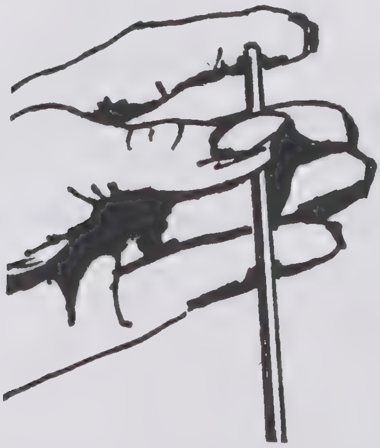


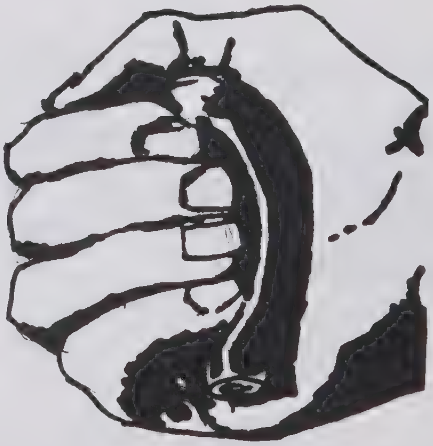
b. graduated pipet

Proper technique for pipetting a sample is:

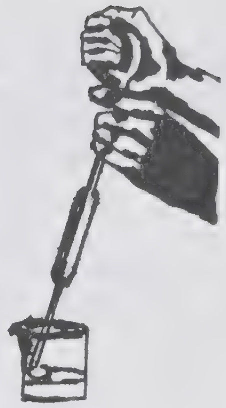
1. Clean and rinse the pipet. (When taking samples of a liquid with a pipet, the pipet is first rinsed with distilled water into a waste beaker following the procedure outlined below. Since the first sample taken is diluted by the film of water already inside the pipet, the first sample is drained into a waste beaker which is known as "rinsing with the sample solution". This procedure should be used whenever a pipet is first used for a given day or whenever a pipet is used for a different solution.)

2. Hold the pipet near the top between the thumb and last three fingers of one hand, leaving the index finger free. Squeeze the pipet bulb closed with the other hand.

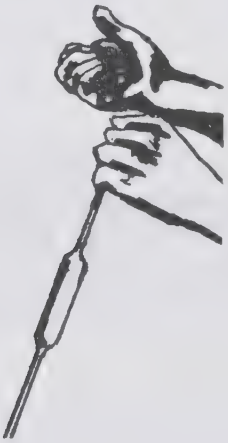




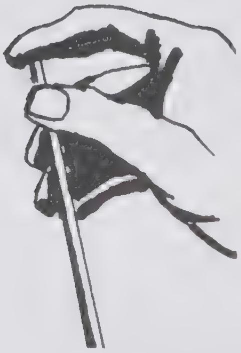
3. Place the tip of the pipet well down into the liquid to be sampled, apply the pipet bulb to the large end of the pipet, and release the pressure on the bulb. This should draw liquid up into the pipet. The bulb should be held against the pipet only firmly enough to make an air seal. Pipetting should NEVER be done using the mouth for suction. Even tiny amounts of some chemicals can be poisonous if taken into the mouth.



4. When the liquid level rises above the top mark on the pipet, remove the bulb and QUICKLY place the index finger over the end of the pipet.

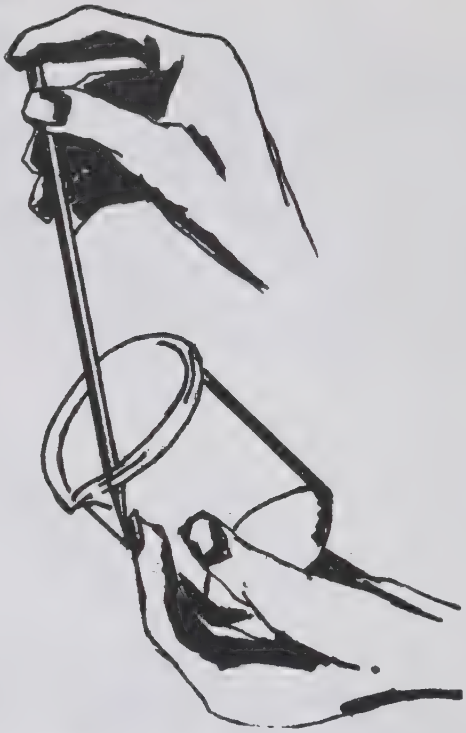


5. Gradually rolling the index finger breaks the air seal and allows liquid to flow out of the pipet. Allow the level to drop until the bottom of the meniscus is exactly on the measuring mark. Hold the level there by pressing harder with the index finger.

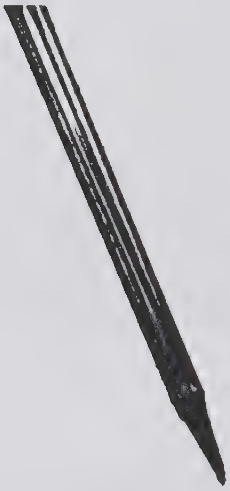


PIPETTING TECHNIQUE - DEMO I2

6. Place the tip of the pipet against the inside wall of the reaction container and allow the contents to flow out of the pipet:
   a. for delivery pipets, simply remove the index finger.
   b. for graduated pipets, stop the flow again when the bottom of the meniscus drops to the graduation mark desired.



7. When a delivery pipe drains, a small amount of liquid remains in the tip. DO NOT BLOW THIS OUT into the reaction container. A delivery pipet is calibrated to correct for this remaining liquid.



SOLUTIONS
PREPARATION OF A STANDARD SOLUTION AND DILUTION OF KNOWN SOLUTION - LAB I1

Purposes:

Part I. To accurately prepare a solution of known concentration.

Part II. To subsequently dilute the solution prepared in Part I.

Apparatus and Materials:

1 - centigram balance
1 - 250 ml beaker
1 - vial of $CuSO_4 \cdot 5H_2O$
1 - wash bottle containing distilled water
1 - stirring rod
1 - funnel
1 - 100 ml graduated cylinder with stopper
1 - 100 ml volumetric flask (if available)
1 - 10 ml graduated cylinder or, preferably, a 10 ml pipet, if available
1 - medicine dropper
1 - meniscus finder
1 - photometer or conductivity apparatus using a multimeter

Procedure:

Part I:

1. Before going to the lab, calculate the number of moles and the mass of $CuSO_4 \cdot 5H_2O$ required to prepare 100.0 ml of 0.2000 M solution. Show your work in the calculation and data section.
2. Determine the mass of a clean, dry 250 ml beaker. Add the desired mass to the riders. Add $CuSO_4 \cdot 5H_2O$ into the beaker until the balance beam levels.
3. Prepare your 0.2000 M $CuSO_4$ solution as follows:
   a. Use your wash bottle to add 40 to 60 ml of distilled water to the $CuSO_4 \cdot 5H_2O$ in the beaker. Stir the solution with a clean stirring rod until the $CuSO_4 \cdot 5H_2O$ is all dissolved. Rinse the stirring rod into the beaker.
   b. Using a clean (not necessarily dry) funnel, transfer the solution into a clean (not necessarily dry) 100 ml graduated cylinder (or 100 ml volumetric flask, if available). Use your wash bottle to rinse the beaker, then the funnel into the graduated cylinder (or volumetric flask).
   c. Use a medicine dropper to carefully bring the bottom of the solution meniscus to the 100.0 ml mark. A meniscus finder is useful here.
   d. Stopper the graduated cylinder (or volumetric flask) and invert several times.

Part II:

4. Dilute your $CuSO_4$ solution as follows:
   a. Pour the 0.2000 M $CuSO_4$ into a clean, dry 250 ml beaker
   b. Transfer 10.0 ml of the 0.2000 M $CuSO_4$ into a clean 100 ml graduated cylinder (or volumetric flask) using a 10 ml graduated cylinder (or preferably a 10 ml pipet, if available). If you are using a 10 ml graduated cylinder, use your wash bottle to rinse the 10 ml graduated cylinder into the 100 ml graduated cylinder (or volumetric flask).

c. Use your wash bottle to carefully add distilled water to the 10.0 ml of $CuSO_4$ solution until the bottom of the solution meniscus reaches the 100.0 ml mark.

d. Stopper the graduated cylinder (or volumetric flask) and invert several times.

e. Take your final solution to your teacher who will check it for colour intensity or conductivity against a set of standard solutions.

Calculations and Data:

1. a. Calculate the number of moles of $CuSO_4 \cdot 5H_2O$ required to make 100.0 ml of 0.2000 M solution.

   b. Calculate the mass of $CuSO_4 \cdot 5H_2O$ required to make the 100.0 ml of 0.2000 M solution.

2. a. Mass of 250 ml beaker ____________

   b. Mass of $CuSO_4 \cdot 5H_2O$ required ____________

   c. Mass of 250 ml beaker plus required $CuSO_4 \cdot 5H_2O$ ____________

3. Calculate the concentration of the solution after dilution.

Questions:

1. What property of the solution changes noticeably upon dilution?

2. Both of the solutions you made would react with zinc metal. Which solution do you predict would react with zinc at a faster rate? Why?

3. Which solution would react with a greater mass of zinc -8.00 ml of the concentrated 0.2000 M $CuSO_4$ solution or 40.0 ml of the diluted solution? Show your calculations.

CHEMISTRY 20
ELECTROLYTES AND NONELECTROLYTES - DEMO I3

Introduction:

Solutes can be classified into two categories:

a. nonelectrolytes which upon dissolving yield solutions that do not conduct electricity, and
b. electrolytes which upon dissolving yield solutions that do conduct electricity.

Purpose:

To classify solutions of compounds as nonelectrolytes or electrolytes.

Procedure:

An electrical conductivity apparatus will be used to determine whether the substances in the following list are nonelectrolytes or electrolytes. Your teacher may wish to add to this list. (All of the solutions are 0.10 M unless otherwise specified.)

- sodium chloride
- hydrochloric acid
- methanol
- sodium hydroxide
- ammonium acetate
- potassium hydroxide
- 1-butanol, C_4H_9OH
- sulfuric acid
- sugar sucrose, $C_{12}H_{22}O_{11}$
- potassium dichromate
- nitric acid
- acetic acid
- ammonium hydroxide
- glucose, $C_6H_{12}O_6$
- calcium hydroxide (saturated)
- iodine (saturated)
- copper(II) sulfate
- potassium permanganate
- acetone, $(CH_3)_2CO$
- glycerol, $C_3H_5(OH)_3$
- sodium bicarbonate

Results:

Summarize the results of the electrical conductivity tests in the following table.

Substances that are Nonelectrolytes	Substances that are Electrolytes

QUESTIONS:

1. What class of compounds were nonelectrolytes? Electrolytes?

2. Propose a hypothesis to explain why some substances are nonelectrolytes while others are electrolytes.

SOLUTIONS

DISSOCIATION

Nonelectrolytes

Nonelectrolytes generally include molecular elements and molecular compounds. Solutions of nonelectrolytes are called nonelectrolytic solutions. When nonelectrolytes dissolve, they separate into individual neutral molecules that are free to move throughout the solution.

Equations showing the dissolving process for nonelectrolytes simply show the solute changing from its pure state to its dissolved state.

For example, when sugar dissolves in water:

$$C_{12}H_{22}O_{11(s)} \longrightarrow C_{1/2}H_{22}O_{11(aq)}$$

Electrolytes

Electrolytes generally include ionic compounds and acids. Solutions of electrolytes are called electrolytic solutions. When electrolytes dissolve, they separate into ions which move freely throughout the solution. When electrolytes dissolve and separate into ions, they are said to dissociate and the process is known as dissociation. Although electrolytic solutions may contain billions of ions, the solutions as a whole are always neutral because the solutions always contain equal quantities of positive and negative charge.

Equations which show the dissolving of electrolytes must show the solute in its pure state changing to aqueous ions. Dissociation equations must:

a. be balanced
b. show correct ionic charges, and
c. show physical states

Examples:

$$KCl_{(s)} \longrightarrow K^{+}_{(aq)} + Cl^{-}_{(aq)}$$

$$Cu(NO_3)_{2(s)} \longrightarrow Cu^{2+}_{(aq)} + 2\ NO_3^{-}{}_{(aq)}$$

$$HBr_{(g)} \longrightarrow H^{+}_{(aq)} + Br^{-}_{(aq)}$$

$$Al_2(SO_4)_{3(s)} \longrightarrow 2\ Al^{3+}_{(aq)} + 3\ SO_4^{2-}{}_{(aq)}$$

DISSOCIATION EQUATIONS

Write dissociation equations for the following electrolytes dissolving and dissociating in water. Show the physical state of all species involved.

1. potassium hydroxide
2. sodium carbonate
3. potassium permanganate
4. hydrogen chloride gas
5. ammonium sulfate
6. gallium carbonate
7. aluminum sulfate
8. sodium bicarbonate
9. potassium nitrate
10. magnesium hydroxide
11. silver nitrate
12. copper (II) nitrate
13. lead (II) nitrate
14. potassium iodide
15. aluminum chlorate

SOLUTIONS

IONIC CONCENTRATIONS IN ELECTROLYTIC SOLUTIONS

Although it is common to refer a "a 0.2 M $Al_2(SO_4)_3$ solution", this is technically incorrect. In the solution the $Al_2(SO_4)_3$ does not exist as $Al_2(SO_4)_3$ units, but exists as Al^{3+} and $SO_4{}^{2-}$ ions. Hence rather than referring to the concentration of the solution, the concentration of the ions is often used.

Calculating Ionic Concentrations

Example 1: In a 0.23 M $Al_2(SO_4)_3$ solution, what is the concentration of each ion?

$$Al_2(SO_4)_{3(s)} \longrightarrow 2\ Al^{3+}{}_{(aq)} + 3\ SO_4{}^{2-}{}_{(aq)}$$

Each $Al_2(SO_4)_3$ unit which dissolves yields two Al^{3+} ions and three $SO_4{}^{2-}$ ions. The concentrations of these ions must be respectively two and three times the concentration of the solute.

$$M_{Al^{3+}} = 0.23M \times 2 = 0.46M$$

$$M_{SO_4{}^{2-}} = 0.23M \times 3 = 0.69M$$

Example 2: In a 0.050 M $(NH_4)_2Cr_2O_7$ solution, what is the concentration of each ion?

$$(NH_4)_2Cr_2O_{7(s)} \longrightarrow 2\ NH_4{}^{+}{}_{(aq)} + Cr_2O_7{}^{2-}{}_{(aq)}$$

$$M_{NH_4{}^+} = 0.050M \times 2 = 0.10M$$

$$M_{Cr_2O_7{}^{2-}} = 0.050M \times 1 = 0.050M$$

Calculating Mass of Solute Needed to Give a Solution of Specified Ionic Concentration

Example: What mass of $Al_2(SO_4)_3$ must be dissolved in water to give a 600 ml solution that has a SO_4 concentration of 0.0500 M?

$$A\ell_2(SO_4)_{3(s)} \longrightarrow 2A\ell^{3+}{}_{(aq)} + 3SO_4{}^{2-}{}_{(aq)}$$

$$mol_{SO_4{}^{2-}} = 0.0500\ \frac{mol}{\ell} \times 0.600\ell = 0.0300\ mol$$

$$mol_{Al_2(SO_4)_3} = 0.0300\ mol \times \frac{1}{3} = 0.0100\ mol$$

$$mass_{Al_2(SO_4)_3} = 0.0100\ mol \times \frac{342\ g}{mol} = 3.42\ g$$

SOLUTIONS
CONCENTRATION OF IONIC SPECIES

For each of the questions (a) to (k):

1. write the dissociation equation, and
2. calculate the concentration of each ion.

a. 0.14 M Na_2SO_4

b. 6.0 M $HC\ell$

c. 2.5 M $(NH_4)_2CO_3$

d. 0.264 M $FeC\ell_3$

e. 0.087 M K_3PO_4

f. 0.126 M $A\ell_2(SO_4)_3$

SOLUTIONS
CONCENTRATION OF IONIC SPECIES

g. 0.329 M barium acetate

h. 1.6 M K_2CO_3

i. 0.85 M lithium sulfide

j. a solution of NaOH containing 2.00g of NaOH in 100 ml of solution.

k. a solution of K_2CrO_4 containing 19.4 g of K_2CrO_4 in 300 ml of solution

SOLUTIONS
CONCENTRATION OF IONIC SPECIES

For each of the questions (ℓ) to (p):

1. write the dissociation equation, and
2. calculate the mass of the compound that must be dissolved to give the ionic concentration specified.

ℓ. Na_2CO_3 to give 1.00ℓ of 0.500 M $CO_3{}^{2-}{}_{(aq)}$ concentration

m. $(NH_4)_2\ SO_4$ to give 200 ml of 1.20 M $NH_4{}^{+}{}_{(aq)}$ concentration

n. $K_2Cr_2O_7$ to give 500 ml of 0.600 M $Cr_2O_7{}^{2-}{}_{(aq)}$ concentration

o. aluminum chlorate to give 150 ml of 1.20 M $C\ell O_3{}^{-}{}_{(aq)}$ concentration

p. rubidium sulfate to give 400 ml of 0.250 M $Rb^{+}{}_{(aq)}$ concentration

SOLUTIONS
SATURATED SOLUTIONS AND SOLUBILITY

Saturated Solutions

Some substances dissolve in each other in all proportions. Examples are water and anti-freeze, and gasoline and kerosene. Such substances are said to be *miscible*. Most other substances are *partially soluble*, that is, one dissolves in the other to a concentration that reaches a *definite limiting value* at which point the solution is said to be *saturated*. For example, when sugar is stirred in water at 20°C, the sugar dissolves rapidly at first, then more and more slowly. Eventually, dissolving appears to stop and the concentration of sugar dissolved in the water no longer increases but remains constant, no matter how long or how vigorously the mixture of sugar and water is stirred (provided the temperature is kept constant). The solution is now *saturated* with sugar. The dissolving action, however, does not cease when the saturation point is reached. Molecules of the sugar continue to leave the solid and pass into the solution, while other molecules of sugar, previously dissolved, return to the solid from the solution. The rate of these two processes is exactly equal at saturation, so that the number of particles of solute leaving the solid and entering the solution in unit time is equal to the number of solute particles leaving the solution and *crystallizing* out on the solid in the same unit of time.

The solution process at saturation is an example of a *dynamic equilibrium: A saturated solution* may be defined as a *solution in which the dissolved solute is in equilibrium with excess undissolved solutes* at a specified temperature.

$$\text{Undissolved solute} \underset{\text{crystallizing}}{\overset{\text{dissolving}}{\rightleftharpoons}} \text{dissolved solute} .$$

Solubility

The term "solubility" is commonly used in two senses - qualitatively and quantitatively. Qualitatively, solubility is often used in a relative way when substances are classed as being soluble, slightly soluble, or very slightly soluble. At extremely low solubilities, the solute may be termed as having negligible solubility or as being insoluble. The table below lists the solubilities generally associated with the qualitative terms.

Solubility	Qualitative Term
greater than 0.1 M	soluble
less than 0.1 M	low solubility
	or in the extreme, solubility is said to be negligible

The qualitative use of solubility is often too imprecise for many purposes. The quantitative definition of solubility has a definite meaning. In the quantitative sense, solubility refers to the quantity of solute required to produce a saturated solution at a given temperature. Solubility is the *concentration of solute in a saturated solution at a given temperature. Thus molar solubility would be the number of moles of solute required to form one litre of saturated solution* at a specified temperature, i.e.,

$$\text{M solute (at saturation)} = \text{solubility.}$$

Example:

Experiment shows that 74.5 g of $CaC\ell_2$ dissolves in water to form 100 ml of saturated solution. Calculate the molar solubility of $CaC\ell_2$.

$$\#\ \text{moles}_{CaC\ell_2} = \frac{\text{mass}}{\text{molar mass}} = \frac{74.5\ \text{g}}{111\ \text{g/mol}} = 0.671\ \text{mol}$$

$$\text{M}_{CaC\ell_2}\ \text{(at saturation)} = \frac{\text{moles}}{\text{litres}} = \frac{0.671\ \text{mol}}{0.100\ \ell} = 6.71\ \text{M}$$

SOLUBILITY TABLE

Name of Ions	Formula of Ions	Soluble With	Low Solubility With
alkali metal	Li^+ etc.	all anions	no anions
hydrogen	H^+	all anions	no anions
ammonium	NH_4^+	all anions	no anions
nitrate	NO_3^-	all cations	no cations
acetate	CH_3COO^-	all cations	no cations
chloride bromide iodide	Cl^- Br^- I^-	nearly all cations	Ag^+, Pb^{2+} Hg^+, Cu^+
sulfate	SO_4^{2-}	nearly all cations	Ag^+, Ca^{2+}, Sr^{2+} Ba^{2+}, Pb^{2+}
sulfide	S^{2-}	all alkali metal ions H^+, NH_4^+ Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	all other cations
hydroxide	OH^-	all alkali metal ions, H^+, NH_4^+, Sr^{2+}, Ba^{2+}	all other cations
phosphate carbonate sulfite	PO_4^{3-} CO_3^{2-} SO_3^{2-}	all alkali metal ions H^+, NH_4^+	all other cations

SOLUTIONS
EXERCISES ON SOLUBILITY

1. 92.1 g of sodium nitrate are dissolved in water to form 100 ml of saturated solution at 25°C. Calculate the molar solubility of sodium nitrate in water at 25°C.

2. Calculate the molar solubility of potassium chloride at 20°C from the following experimental data:
   mass of evaporating dish . . . . = 43.51 g
   volume of $KC\ell$ saturated solution placed in evaporating dish . . . . = 20.0 ml
   mass of evaporating dish and dry residue after evaporation of water. . = 50.45 g

3. An important component of plaster of Paris and of gypsum, $CaSO_4$, has a solubility of 0.0239 moles per litre at 25°C. What mass of $CaSO_4$ would remain if 200 ml of this solution were evaporated to dryness?

4. The molar solubility of sugar in water is 3.80 M at room temperature. Assuming that the mass of a sugar cube is 3.80 g and that the volume of an ordinary cup is 250 ml, calculate the number of sugar cubes that would be required to prepare a sugar saturated cup of water.

SOLUTIONS
FACTORS THAT AFFECT SOLUBILITY

The prediction of solubilities involves many variables such as relative size and relative charge of solute and solvent particles, interaction between solute and solvent particles, temperature, and pressure. In spite of the complexity of factors that affect solubility, some general rules apply to many of the compounds first encountered in the study of chemistry. These general rules are not laws and are therefore subject to exceptions. A discussion of the four factors affecting solubility and some relevant general rules follows.

1. The Nature of Solute and Solvent

The rule "like dissolves like" where "like" refers to similarities in polarities of substances, has useful application for predicting solubilities. In general, polar and ionic solutes tend to be more soluble in polar solvents and nonpolar solutes tend to be more soluble in nonpolar solvents. Thus, inorganic acids, which are polar and bases and salts, which are ionic, tend to be much more soluble in water, which is highly polar. Solvents such as carbon tetrachloride, hexane and benzene, which are nonpolar, are not soluble in polar water. Sodium chloride, an ionic substance, is highly soluble in polar water, slightly soluble in weakly polar ethyl alcohol and insoluble in nonpolar carbon tetrachloride. Gasoline, a nonpolar substance, is only very slightly soluble in polar water but highly soluble in nonpolar carbon tetrachloride.

General Rules:

a) Polar and ionic solutes are soluble in polar solvents.

b) Nonpolar solutes are soluble in nonpolar solvents.

2. The Effect of Temperature

Since most solutes have a limited solubility in a given amount of solvent at a fixed temperature, the temperature of the solvent generally has a marked effect on the amount of solute that will dissolve. For most *solids* dissolved in liquids, the dissolving process is endothermic and *an increase in temperature results in an increase in solubility*. The effect of increased solubility of solids in liquids upon heating is illustrated in everyday experiences. For example, the solubility of sugar in heated tea or coffee is greater than in cold tea or coffee. For most *gases*, the dissolving process is exothermic and *a decrease in temperature results in an increase in solubility*. For example, the hissing noise and the bubbles which form when water in a kettle is heated but not boiled are due to the fact that dissolved air escapes as it becomes less soluble at higher temperatures. For the dissolving of liquids in liquids, the effect of temperature is too variable and no useful generalization can be made. In every case, a new solubility under new temperature conditions is gradually established.

General Rules:

a) An increase in temperature generally increases the solubility of most solids in liquids.

b) An increase in temperature generally decreases the solubility of most gases in liquids.

3. The Effect of Pressure

Changes in pressure have very little effect on the solubility of solids and liquids. However, changes in pressure have a marked effect on the solubility of gases. The solubility of a gas in a liquid is directly proportioned to the pressure of that gas above the liquid. For example, when a bottle of carbonated soft-drink is opened, the pressure is reduced and dissolved carbon dioxide bubbles out of thesolution. A new solubility under the new pressure conditions is gradually established.

General Rules:

1. Changes in pressure have no appreciable effect upon the solubility of solids and liquids.

2. The solubility of gases in liquids is directly proportional to the pressure of the specific gas above the liquid.

I9, 10

THE SOLUBILITY OF A SOLUBLE SUBSTANCE - LAB I2

For this experiment each group will follow a slightly different procedure. At the end of the lab results should be shared with other groups. Each group will prepare a saturated solution at either of three temperatures - at the temperature of melting ice, at room temperature and at the temperature of boiling water. Your teacher will assign you the temperature for your particular solution.

Purpose:

1. To determine the solubility of Na_2CO_3 in water.
2. To compare the solubility of Na_2CO_3 at different temperatures.

Apparatus and Materials:

1 - centigram balance
1 - 100 ml graduated cylinder
1 - 400 ml beaker
1 - 100 ml beaker
1 - buret clamp
1 - ring stand
1 - wash bottle containing distilled (deionized) water
1 - 10 ml pipet (or 10 ml graduated cylinder)
1 - petri dish
1 - stirring rod
1 - thermometer
- about 20 g Na_2CO_3

Procedure:

1. Obtain about 20 g of $Na_2CO_{3(s)}$ and place it into a 100 ml beaker.
2. Add about 40 ml of distilled water to the $Na_2CO_{3(s)}$ in the 100 ml beaker.
3. (a) If your experiment is to be done at *room temperature* continue with procedure step 5.
   (b) If your experiment is to be done at the *temperature of melting ice*, prepare a water and ice mixture in a 400 ml beaker about three-fourths full. (More effective cooling can be obtained by adding some $NaC\ell$ to the ice-water mixture.) Support the 100 ml beaker containing the Na_2CO_3 solution using a buret clamp and ring stand. Lower the Na_2CO_3 solution into the ice-water mixture so that the level of the solution is below the level of water in the ice-water. Continue with procedure step 4.
   (c) If your experiment is to be done at the *temperature of boiling water*, fill a 400 ml beaker about three-fourths full and heat this water to boiling. Support the 100 ml beaker using a buret clamp and ring stand. Lower the Na_2CO_3 solution into the ice-water mixture so that the level of the solution is below the level of the boiling water. Continue with procedure step 4.
4. Allow several minutes for the Na_2CO_3 solution to attain the temperature of the ice-water mixture or the boiling water.
5. Stir the solution until no more sodium carbonate dissolves, i.e., until equilibrium between the solution and solute is attained.
6. Label a petri dish with your initials.
7. Determine the mass of the clean, dry petri dish to the nearest 0.01 g.
8. Determine the temperature of the saturated Na_2CO_3 solution.
9. Using a 10 ml pipet (or graduated cylinder) transfer 10 ml of the saturated solution from the beaker to the petri dish. Be careful NOT to transfer any Na_2CO_3 solid.

10. Allow the solution to evaporate. (Your teacher will indicate a method for evaporation of the solution.)

11. When the residual Na_2CO_3 is dry, redetermine the mass of the petri dish and Na_2CO_3 to the nearest 0.01 g.

12. Do the required calculations. Summarize the class results for molar solubility of Na_2CO_3 at the three different temperatures. Your teacher will suggest a convenient method for summarizing the class results.

Data:

Mass of petri dish . . . . ____________________

Mass of petri dish + $Na_2CO_{3(s)}$ . . . . ____________________

Temperature of saturated Na_2CO_3 solution . . . . . ____________________

Calculations: (show method of calculation.)

1. Mass of crystallized Na_2CO_3 in petri dish.

2. Number of moles of crystallized Na_2CO_3 in petri dish.

3. Molar concentration of saturated Na_2CO_3 solution.

Summary of Class Data:

Temperature	Molar Solubility

QUESTIONS:

1. Explain the relationship between a saturated solution and solubility.

THE SOLUBILITY OF A SOLUBLE SUBSTANCE - LAB I2

2. According to the data obtained, what is the effect of temperature upon the molar solubility of Na_2CO_3?

3. Na_2CO_3 is an electrolyte and upon dissolving in water dissociates into its component ions.

   a) Write an equation to represent the dissociation of Na_2CO_3 in water.

   b) Keeping in mind the balanced equation representing the dissociation of Na_2CO_3 in water, give the molar concentration of $Na^+_{(aq)}$ and $CO_3{}^{-2}_{(aq)}$ in a saturated Na_2CO_3 solution at room temperature.

4. Discuss the statement, "molar solubility gives the concentration at saturation of a solution of any volume".

DATE DUE SLIP

JUL 27 RETURN
DUE EDUC MAR 5 '81
RETURN MAR 4 - '81
DUE EDUC NOV 25 '82
RETURN NOV 20 '82
DUE EDUC MAR 13 '83
RETURN MAR 9 '83
DUE EDUC. SEP 22 '92
1992 SEP 22 RETURN
DUE EDUC. OCT 16 '92
1992 OCT 13 RETURN

F255

JUL 7 1977

QD 41 A3312 1976 UNIT I C-3
ALBERTA CHEMISTRY PROJECT
MATERIALS
ALCHEM 20
39283345 CURR

000020716080

RECOMMENDED FOR USE
IN ALBERTA SCHOOLS

QD

QD 41 A3312 1976 Unit I C. 3
Alberta Chemistry Project
Materials.
ALCHEM 20.
0187415B CURR

C. 3

CURRICULUM
EDUCATION LIBRARY

B15943